CHAPTER 7 – Periodic Properties of the Elements

Section 7.1 – Development of the Periodic Table

(a) In 1869, Dmitri Mendeleev and Lothar Meyer organized their periodic tables on the basis of increasing **atomic mass**. Why didn't these scientists choose to arrange the elements in order of increasing **atomic number**?

(b) Mendeleev is generally given credit for discovering the periodic table, even though both he and Meyer had very similar arrangements. The likely reason for this is that Mendeleev had the courage to leave ________________________________ in his table to make room for elements that he predicted should exist, but had not yet been discovered.

(c) What two elements were called *eka-aluminum* and *eka-silicon* by Mendeleev?

________________ ________________

(d) In 1913, Moseley's work developed the concept of ___________________________________

(e) The average atomic mass of iodine (126.90) is less than the average atomic mass of tellurium (127.60). Explain why iodine is placed to the right of tellurium on the periodic table.

Section 7.2 – Effective Nuclear Charge

Go to page 46 in your textbook. ("A Closer Look: Basic Forces") and read the information about Coulomb's law.

(a) Coulomb's law describes the magnitude of the ____________________________ between two charged particles.

(b) The force between two charged particles is _________________ proportional to the magnitude of the charges (Q_1 and Q_2).

(c) The force between two charged particles is _________________ proportional to the square of the distance between them (d^2).

(d) If the two charges are of opposite sign, the force between them is ______________________.

If the charges have the same sign, the force is ____________________________.

(e) The attractive force between an electron and the ____________________ depends on the magnitude of the nuclear charge and the average distance between them.

(f) As nuclear charge increases, the attractive force ____________________

(g) As the electron moves farther from the nucleus, the attractive force ______________________

(h) The net electric field created by the nucleus and the electron density of the other electrons is called the ____, symbolized as Z_{eff} .

(i) In any many-electron atom, the inner electrons (core electrons) partially screen outer electrons from the ____.

(j) The effective nuclear charge _______________________ from left to right across any period of the periodic table.

Going down a column, the effective nuclear charge experienced by valence electrons changes far less than it does across a period.

<u>Section 7.3 – Sizes of Atoms and Ions</u>

(a) We can define the ____________________________ as half of the distance between the two nuclei of bonded atoms.

(b) Within each group on the periodic table, the atomic radius tends to ______________________ from top to bottom as you travel down the group. This trend results primarily from the ____________________ in the principal quantum number (*n*) of the outer electrons as you go down a column.

(c) Within each period on the periodic table, atomic radius tends to _________________________ from left to right. The main reason for this trend is the _________________________________ ____ across a period.

(d) Based on their relative positions on the periodic table (and not looking at the actual values for atomic radius shown in Figure 7.6), decide which element has the larger atomic radius.

Elements to be compared	Which element should have a larger atomic radius?
Na and K	
Al and Si	
S and F	
O and Cl	

(e) When a cation is formed from a neutral atom, electrons are ____________________ from the valence shell. The ionic radius of a cation will always be ____________________ than the atomic radius of its parent atom.

(f) When an anion is formed from a neutral atom, electrons are ____________________ to the valence shell. Because of the increase in electron-electron repulsions, the ionic radius of an anion will always be ______________________ than the atomic radius of its parent atom.

(g) When comparing ions that have the same charge in the same group, the ionic radius ______________________ as we move from top to bottom. This is because the principal quantum number (*n*) of the outermost occupied orbital of the ion is increasing.

(h) A group of ions that all contain the same number of electrons is called an _________________________ series. If we compare two or more ions that have the same number of electrons, the ion that has the ____ should have the smallest ionic radius.

(i) Arrange these five ions in order from smallest to largest radius: O^{2-}, S^{2-}, F^{-}, Na^{+}, Mg^{2+}

__________ __________ __________ __________ __________

smallest largest

<u>Section 7.4 – Ionization Energy</u>

(a) Define the term ionization energy.

(b) The value of ionization energy is often measured in kJ/mol. The greater the ionization energy, the ______________________ it is to remove an electron.

(c) Examine the data in Table 7.2 on page 259. Why do the values for ionization energies for a given element increase as successive electrons are removed from the atom?

(d) After all of the valence electrons of an atom have been removed, there is a sharp increase in the value for ionization energy for the next electron that is to be removed from the atom. Explain why this sharp increase in I.E. occurs.

(e) The inner core electrons are very tightly bound. Therefore when electrons are shared and/or transferred in chemical reactions, only the ______________________ electrons are involved.

(f) The 1st I.E. generally ___________________ as we move from left to right across a period. As we move across a period, the effective nuclear charge ___________________ and the atomic radius __________________. Thus the electrons are more attracted toward the nucleus of the atom.

(g) The 1st I.E. generally ___________________ as we go from top to bottom down a group. As we move down a column, the atomic radius __________________ while the effective nuclear charge increases rather gradually.

(h) In general, the smaller the atomic radius, the ___________________ the 1st I.E. will be. The energy required to remove an electron from the outermost shell depends on both the _____________________________________ and the average __________________________ ____.

(i) Explain why the 1st I.E. for boron (801 kJ/mol) is slightly less than the 1st I.E. for beryllium (899 kJ/mol)

(j) Explain why the 1st I.E. for oxygen (1314 kJ/mol) is slightly less than the 1st I.E. for nitrogen (1402 kJ/mol)

(k) When electrons are removed from an atom to form a cation, they are always removed from the orbital that has the ________________________ principal quantum number. For example, the electron that is lost from lithium comes from the ____ orbital.

(l) When an atom of iron loses 2 electrons to form a +2 ion, the electrons are removed from the __________ orbital. If another electron is removed to form a +3 ion, this electron is removed from the _________ orbital.

(m) If there is more than one occupied subshell for a given value of *n*, the electrons are first removed from the orbital that has the ____

(n) Electrons added to an atom to form an anion are added to the empty or partially filled orbital that has the ____

Section 7.5 – Electron Affinities

(a) Is removing an electron from an atom (ionization) an endothermic process or an exothermic process? Explain.

(b) Define electron affinity.

(c) When a chlorine atom acquires an electron to form the chloride ion, energy is released (–349 kJ/mol). What does this information tell us about the relative stability of the Cl atom and the Cl ion?

(d) Why is the electron affinity for noble gas atoms a positive number?

Section 7.6 – Metals, Nonmetals, and Metalloids

(a) Metallic character generally __________________ as we proceed down a group of the periodic table and it __________________ as we move from left to right across a period.

(b) Describe the physical characteristics that tend to be common among most metals.

(c) Metals tend to have ______________ ionization energies, and during chemical reactions they tend to become (oxidized reduced).

(d) What fundamental property of an element is the best indicator of whether it behaves as a metal or a nonmetal ? ____

(e) Metals and nonmetals tend to react together form ______________ substances. Most metal oxides behave as ______________, which, if they are soluble, form metal hydroxides in aqueous solution. Insoluble metal oxides should be soluble in ____________ solutions, because they will react to form water and a salt.

(f) Write a balanced chemical equation for each of the following.

Solid potassium oxide reacts with water.

Solid copper(II) oxide reacts with aqueous sulfuric acid.

(g) Describe the physical characteristics that tend to be common among most nonmetals.

(h) When nonmetals react with metals, which type of element normally gains electrons and which type normally loses electrons?

(i) If a nonmetal oxide is soluble in water, it reacts to form an ______________ solution. When nonmetal oxides react with bases, they tend to form _____________ and ______________.

(j) Write a balanced chemical equation for each of the following.

Sulfur trioxide reacts with water.

Diphosphorus pentoxide reacts with aqueous sodium hydroxide.

(k) Metalloids, such as _________________ have properties intermediate between metals and nonmetals, which is why they can be used in semiconductors.

Section 7.7 – Trends for Group 1A and Group 2A Metals

(a) Alkali metals tend to have ________ densities and ________ melting points.

(b) The alkali metal of any given period has the __________ value for 1^{st} I.E. in that period.

(c) Write a balanced chemical equation for each of the following.

Lithium reacts with hydrogen.

Potassium reacts with water.

(d) Compared with alkali metals, alkaline earth metals have ______________ densities and _________________ melting points.

(e) Examine each of the following reactions. Would they occur or not? If so, write the products and balance the equation.

$Sr + H_2O \rightarrow$

$Mg + Ba(NO_3)_2 \rightarrow$

$Ba + Mg(NO_3)_2 \rightarrow$

$Ca + N_2 \rightarrow$

Section 7.8 – Trends for Selected Nonmetals

(a) How does the ionization energy of hydrogen compare with the other elements in Group 1?

(b) Write the products for each reaction and balance the equation.

$H_2 + K \rightarrow$

$H_2 + Cl_2 \rightarrow$

$H_2 + N_2 \rightarrow$

$Ca + HNO_3 \rightarrow$

(c) What members of Group 16 (Group 6A) are solids at room temperature?

(d) What are the two allotropes of oxygen, and which one is more stable?

(e) As the atoms in Group 16 become larger in size, they become ____________ metallic in their properties.

(f) Identify the state of matter present at room temperature for each of these halogens:

F_2 ___________ Cl_2 __________ Br_2 __________ I_2 ____________

(g) Would the following reaction occur? If so, what are the products? If not, why not?

$Br_2 + NaF \rightarrow$

(h) The noble gas of any given period has the ____________ value for 1st I.E. in that period.

(i) What are the possible products of the following reaction?

$Xe + F_2 \rightarrow$